Unknown # 25

Objective: To titrate an unknown sample of acid with a standardized base solution and determine the CONCENTRATION of the unknown in Molarity.

Reaction: $NaOH_{(aq)} + CH_3COOH_{(aq)} \rightarrow NaCH_3COO_{(aq)} + H_2O_{(\ell)}$

PROCEEDURE:

Observations:
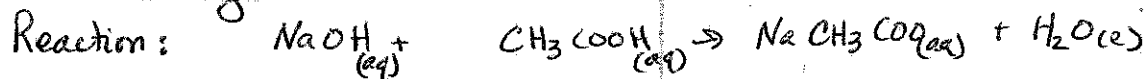
1. Obtain a solution of Standerdized base.

   Concentration NaOH   0.1023 M

2. Prime a burett, ~~by~~ with ~10 mL of Base

3. Aquire a sample and record unknown number (also at top!)

   unknown number   25

   Better: also record observations. Solution colors, smell, etc? Especially if changes occur!

4. Quantitatively transfer unknown to 100mL volumetric flask and fill to the line with distilled water.

5. ~~USE A pipet to Transfer~~ prime a ~~10.00~~ 15.00mL mL pipet with the unknown Acid solution

   Changed per prelab instruction

   Note changes to procedure as needed.

6. pipet ~~10.00mL~~ 15.00mL of Acid into an 125 mL Erlenmeyer flask, Add ~2 drops phenopthalene indicator, and ~25mL distilled water.

   Trial data for each trial in tablel on next page.

7. Titrate the erlenmyer flask mixture with the base in the buret until the solution is very light pink

8. Repeat until reproducible (4 or more trials).

Data:

Table 1: Titration Trials

Record data, be clear.

Trial #	T1	T2	T3	T4	T5
Amount of Acid (mL)	~~10.00~~ 15.00	15.00	15.00	15.00	15.00
Initial buret reading (mL)	0.00	13.24	0.05	12.06	24.06
Final buret reading (mL)	13.24	25.74	12.06	24.06	36.09
Volume base (mL)	13.24	12.50	12.01	12.00	12.03 mL
~~Molarity of Unknown Acid (M)~~					
Endpoint Color	dark pink	dark pink	light pink	light pk	light pink
Valid or Not, why.	No past endpt	No	good	good	good
Molarity of unknown Acid (M)			0.08225M	0.08184	0.08204

> Buret dripped to fast!

Calculations

T3 Molarity of unknown Acid

$$\frac{0.1023 \text{ mol NaOH}}{\text{L Solution}} \times \frac{1 \text{ L Sol'n}}{1000 \text{ mL Sol'n}} \times \frac{12.06 \text{ mL}}{\text{Sol'n}} \times \frac{1 \text{ mol Acetic Acid (AA)}}{1 \text{ mol NaOH}} \times \frac{1}{15.00 \text{ mL AA. Sol'n}} \times \frac{1000 \text{ mL}}{1 \text{ L}} =$$

$$= 0.0822492 \frac{\text{mol AA}}{\text{L Sol'n}} \Rightarrow \boxed{0.08225 \text{ M AA}}$$

Shown calculations are clear with units and easy to follow.

You can also have calculation result in tables. Replicate trials may be calculated using excel, but show calculation in notebook also to chec excel formula results.

Trials 3 → 4 standard deviations

T1 0.08225
T2 0.08184
T3 0.08204

$\bar{x} = 0.082043 \Rightarrow 0.08204 = \text{Average}$

Std deviation

$$S = \sqrt{\frac{\Sigma(x-\bar{x})^2}{n-1}}$$

$= 0.0002050$

$$\% RSD = \frac{0.0002050}{0.08204} \times 100 = \boxed{0.2487\%}$$

These calculations completed in Microsoft Excell.

You don't have to show average or standard deviation calculations as I assume you can do them correctly However you must be clear about what trials are being used and why

Calculations Continued.

95% Confidence Interval.

$$\bar{X} \pm \frac{t \cdot s}{\sqrt{n}} \Rightarrow 0.08204 M \pm \frac{4.303 \cdot 0.0002050}{\sqrt{3}} \Rightarrow \boxed{0.08204 M \pm 0.00051 M}$$

t from table of students t values for 95% Confidence from text book.

Results.

Table 1 contains the molarity $\overset{\text{of the unknown Acid}}{\vee}$ for Trials 3, 4, 5. The average molarity was found to be 0.08204 M with a standard deviation of 0.0002050 M. The 95% confidence interval for these three trials was 0.08204 M ± 0.00051 M.

Conclusions.

For Sample # 25 the Average molarity was found to be 0.08204 M. Acetic Acid with a standard deviation of 0.0002050 M over 3 trials which has a 95% confidence limit of 0.08204 M ± 0.00051 M.

Discussion:

In this experiment, a standardized solution of NaOH was used to titrate an unknown solution of acetic acid. The 95% confidence interval of 0.08204 M ± 0.00051 M says that with 95% confidence, the real CONCENTRATION of the acid is between 0.08153 M and 0.08255 M.

One possible error that could have occured was with the given CONCENTRATION of the NaOH, if this was incorrect the resulting molarity found would also be correct. a better method would be to experimentally standardize the solution with a primary standard acid.
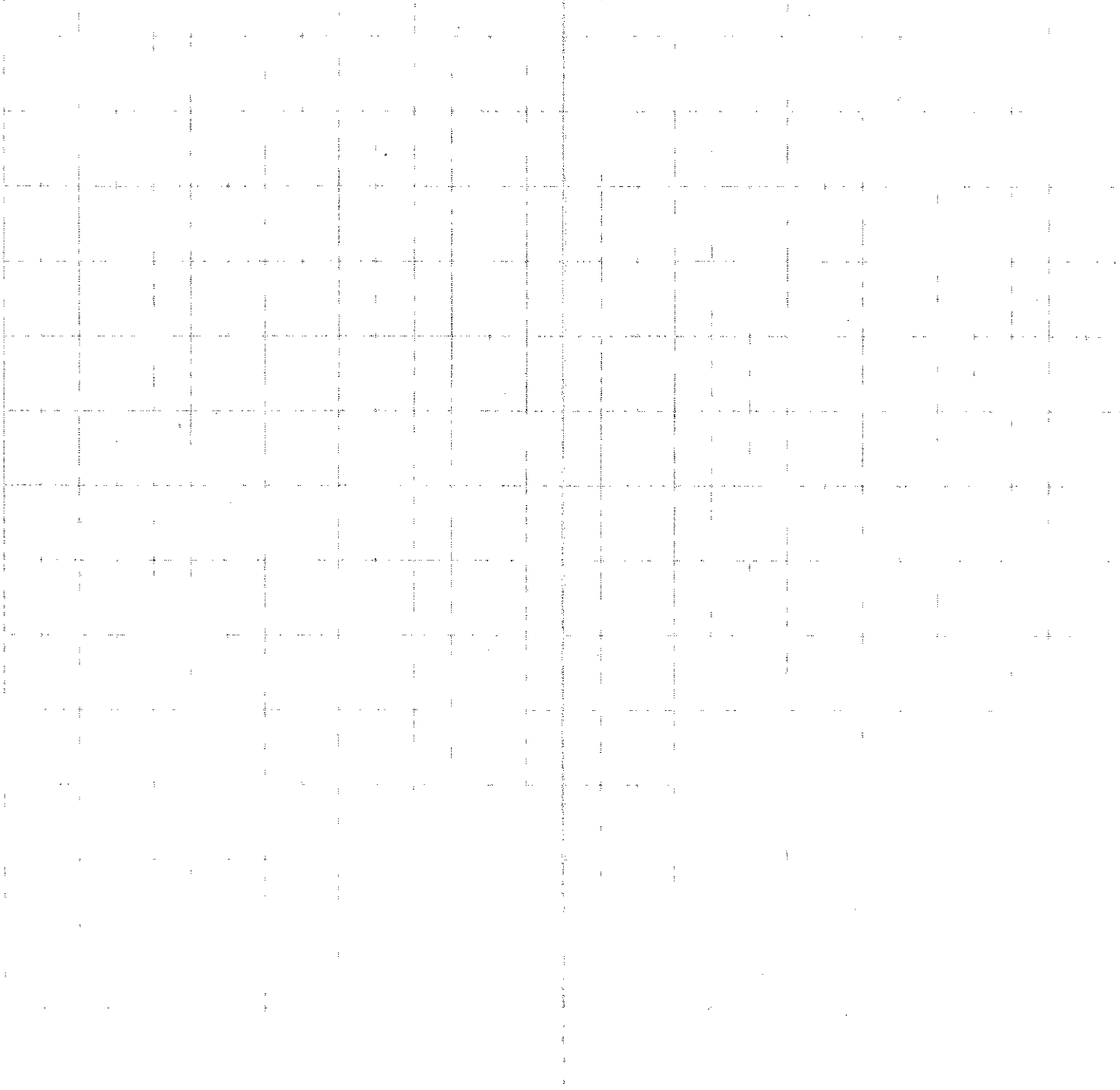
The second source of error could be the 15.00 mL pipet and/or the buret volume measurements are very important for both the unknown acid and standardized base.

Discussion Cont'd.

A third source of error was the indicator used. If the indicator changed colors after the ~~end point~~, actual equivalence point, the calculated molarity of the Acid would be too high. This is why the proper selection of an Indicator is key.